The Mechanism of the Reaction of Phenylpropargylaldehyde with Aqueous Sodium Hydroxide to Give Phenylacetylene and Sodium Formate[1a]

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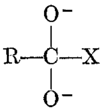
The rate of cleavage of phenylpropargylaldehyde in the presence of aqueous sodium hydroxide at 35.2° has been studied spectrophotometrically. The results were interpreted in terms of rate-controlling cleavage of the monoanion and the dianion of the aldehyde hydrate, with the latter being the principal reaction path at the base concentrations (>0.001 *M*) used. An approximate value was obtained for the equilibrium constant for hydration of the aldehyde and from it was obtained an approximate ionization constant for the aldehyde hydrate, which was found to fit a Taft equation correlation of the acidities of aldehyde hydrates.

Kinetic studies have shown that the cleavage of the carbanion X^- from compounds of the type RCOX by the action of hydroxide ions involves, at least in part, two hydroxide ions in cases where X is trichloromethyl,[2,3] pyridinomethyl,[4] acetonyl,[5] 1-acetylethyl,[5] α-arylsulfonylmethyl,[6] and 2,6-dihalophenyl.[7] These observations may be explained by a mechanism involving rate-controlling loss of X^- from the dianion **1**.



$$\mathrm{R{-}\overset{\overset{\large O^-}{|}}{\underset{\underset{\large O^-}{|}}{C}}{-}X}$$

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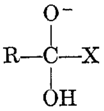

$$\mathrm{R{-}\overset{\overset{\large O^-}{|}}{\underset{\underset{\large OH}{|}}{C}}{-}X}$$

2

Of the compounds undergoing cleavage *via* the dianion, only in the cases where X was acetonyl or 1-acetylethyl was the reaction involving only one hydroxide ion also clearly detectable under the alkaline conditions used.[8] In these cases a significant fraction of the reaction presumably consists of the rate-controlling loss of X^- from the monoanion **2**. Cleavage *via* the monoanion appears to take place in cases in which X is a 1,1-bis(carbalkoxy)alkyl group[9] or a cyano group[10] in acidic solutions, where the concentration and rate of formation of the dianion must be very small, but in more basic solutions the rate-controlling step becomes attack on the carbonyl group. Kinetic studies show the participation of only one hydroxide ion in alkaline cleavages where X is 2-acetyl-2-propyl[5] and nitromethyl.[6] Although it is likely that loss of X^- from anion **2** is rate-controlling in these cases (especially the former), this has not been established. The cleavage of the hydrate of a 2-acetylthiazolium ion involves rate-controlling cleavage of ion **2** in the pH range of 5–7, but the reaction became too fast to study in basic solutions.[11] In order to learn more about the factors that control the relative extent to which X^- is lost from ions **1** and **2**, we have studied the kinetics of the cleavage of phenylpropargylaldehyde, which was found by Claisen to yield phenylacetylene and formate in the presence of dilute aqueous alkali.[12]

Results and Discussion

The cleavage of phenylpropargylaldehyde in 6.5×10^{-5} *M* aqueous solutions was studied kinetically in the presence of about 0.001–0.014 *M* sodium hydroxide by following the decrease in absorbance at the 285.5-nm aldehyde absorption maximum. The absorption spectrum of a reaction solution after about 11 half-lives was in good agreement with that of phenylacetylene formed in 96% yield. Although phenylacetylene and sodium formate do not absorb appreciably at 285.5 nm and at the time of measurement less than 0.05% of the aldehyde should be left, there was an "infinity" absorbance equal to 3.6% of the initial absorbance. Plots of $\ln (A_t - A_\infty)$ *vs.* time (where A_t is the absorbance at time t and A_∞ that at infinite time) gave good straight lines from whose slopes first-order rate constants were calculated. These rate constants will be treated in terms of mechanism 1, where the abbreviations to be used for various species are written under their formulas. There is evidence that

$$\underset{\mathrm{P}}{\mathrm{PhC{\equiv}CCHO}} + \mathrm{H_2O} \overset{K_H}{\rightleftharpoons} \underset{\mathrm{POH_2}}{\mathrm{PhC{\equiv}CCH(OH)_2}}$$

$$\mathrm{PhC{\equiv}CCH(OH)_2} + \mathrm{OH^-} \overset{K_1}{\rightleftharpoons} \mathrm{PhC{\equiv}CCH(OH)O^-}$$

$$\underset{\mathrm{POH^-}}{\mathrm{PhC{\equiv}CCH(OH)O^-}} + \mathrm{OH^-} \overset{K_2}{\rightleftharpoons} \underset{\mathrm{PO^{2-}}}{\mathrm{PhC{\equiv}CCH(O^-)_2}} \qquad (1)$$

$$\mathrm{PhC{\equiv}CCH(OH)O^-} \xrightarrow{k_1} \mathrm{PhC{\equiv}C^-} + \mathrm{HCO_2H} \xrightarrow{\text{fast}} \mathrm{PhC{\equiv}CH} + \mathrm{HCO_2^-}$$

$$\mathrm{PhC{\equiv}CCH(O^-)_2} \xrightarrow{k_2} \mathrm{PhC{\equiv}C^-} + \mathrm{HCO_2^-} \xrightarrow[\mathrm{H_2O}]{\text{fast}} \mathrm{PhC{\equiv}CH} + \mathrm{HCO_2^-}$$

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(2) C. Gustafsson and M. Johanson, *Acta Chem. Scand.*, **2**, 42 (1948).

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(4) R. G. Pearson and R. L. Dillon, *J. Amer. Chem. Soc.*, **70**, 1933 (1948).

(5) R. G. Pearson and E. A. Mayerle, *ibid.*, **73**, 926 (1951).

(6) R. G. Pearson, D. H. Anderson, and L. L. Alt, *ibid.*, **77**, 527 (1955).

(7) J. F. Bunnett, J. H. Miles, and K. V. Nahabedian, *ibid.*, **83**, 2512 (1961).

(8) A reaction involving only one hydroxide ion was reported in the case where X was trichloromethyl,[2] but under the conditions where this reaction is most important the formation of chloride ions becomes a major side reaction[3] that was not taken into account.

(9) G. E. Lienhard and W. P. Jencks, *J. Amer. Chem. Soc.*, **87**, 3855 (1965).

(10) F. Hibbert and D. P. N. Satchell, *J. Chem. Soc. B*, 653 (1967).

(11) G. E. Lienhard, *J. Amer. Chem. Soc.*, **88**, 5642 (1966).

(12) L. Claisen, *Ber.*, **31**, 1022 (1898).

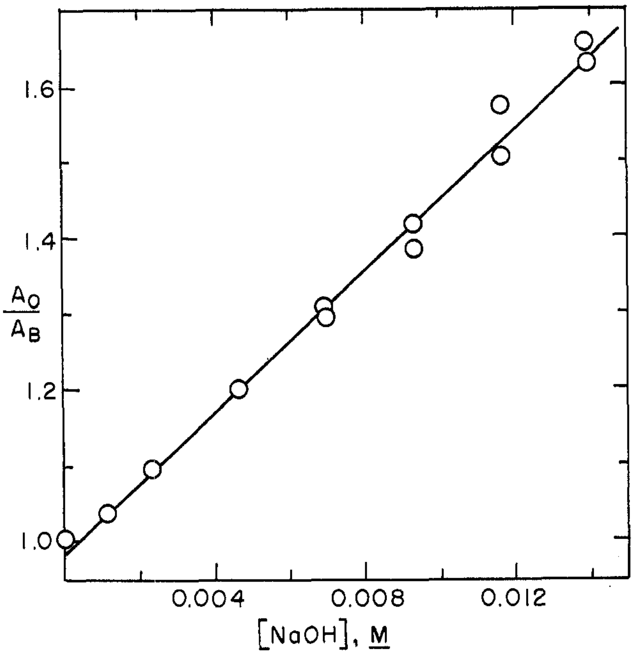

Figure 1.—Plot of ratio of absorbance in the absence of base to the absorbance in the presence of base *vs.* base concentration for phenylpropargylaldehyde in water at 35.2°.

significant fractions of the aldehyde are present in each of the three forms, P, POH_2, and POH^-, but K_2 is believed to be so small that the concentration of PO^{2-} is much lower. If this assumption is correct, and if the rate-controlling steps of the reaction are those governed by k_1 and k_2, the observed first-order rate constant for the reaction at a given hydroxide ion concentration may be expressed as shown in eq 2,

$$k_{\text{obsd}} = \frac{K_E[OH^-](k_1 + k_2K_2[OH^-])}{1 + K_E[OH^-]} \tag{2}$$

where K_E is defined as $(K_1K_H)/(1 + K_H)$. The hydration equilibrium of the aldehyde was found to be established almost as rapidly as absorbance measurements could be made after preparation of a solution in neutral water. It therefore appears safe to assume that the equilibria governed by K_H and K_1 are established before measurements can be made on mixtures of aqueous aldehyde and alkali solutions. This and the assumption that the aldehyde hydrate is too weakly acidic to ionize appreciably in neutral solution lead to eq 3, in which A_0 is the absorbance of a neutral

$$A_0/A_B = 1 + K_E[OH^-] \tag{3}$$

solution of aldehyde and A_B is the absorbance, extrapolated to zero time, of a solution containing the same total amount of aldehyde and the indicated concentration of hydroxide ions. Values of A_0 and A_B are listed with the kinetic data in Table I. A plot of A_0/A_B *vs.* $[OH^-]$,[13] as shown in Figure 1, gives a satisfactory straight line with the intercept 0.98, in reasonable agreement with the value required by eq 3, and slope (K_E) 46.5 M^{-1}. Rearrangement of eq 2 gives eq 4.

$$\frac{k_{\text{obsd}}(1 + K_E[OH^-])}{K_E[OH^-]} = k_1 + k_2K_2[OH^-] \tag{4}$$

Using the value of K_E obtained from eq 3, the left-hand side of eq 4 was plotted against the hydroxide ion concentration[14] (Figure 2), to give a satisfactory straight line whose intercept (k_1) is 9.3 × 10^{-5} sec^{-1} and whose slope (k_2/K_2) is 0.299 M^{-1} sec^{-1}. Since the intercept is fairly near the origin, the value of k_1 is probably much less reliable than that of k_2K_2. From these values it may be calculated that, even in the most favorable case, less than

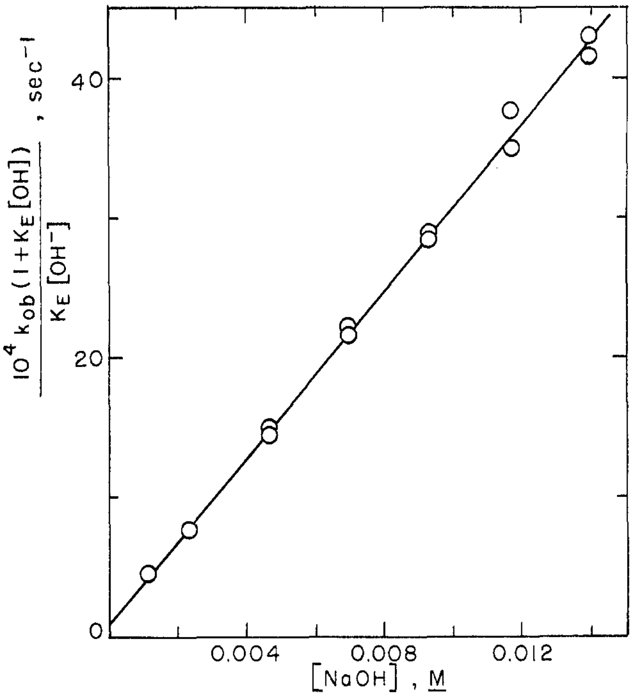

Figure 2.—Plot of the left-hand side of eq 4 *vs.* the sodium hydroxide concentration.

TABLE I

KINETICS OF ALKALINE CLEAVAGE OF PHENYLPROPARGYLALDEHYDE IN WATER AT 35.2° [a]

10^3 [NaOH],[b] M	A_B[c]	10^5 k_{obsd}, sec^{-1}
4.67	0.535[d]	26.5
7.00	0.496[d]	54.4
9.34	0.463[d]	87.2
11.7	0.426[d]	123.0
14.0	0.394[d]	164.2
1.16	0.620	2.34
2.33	0.586	7.41
4.66	0.539	25.7
6.99	0.493	53.0
9.32	0.454	85.9
11.6	0.409	132.2
14.0	0.389	169.1
11.6[e]	0.415	151.8
11.6[f]	0.418	115.0

[a] The ionic strength was adjusted with sodium chloride and was 0.10 M except where otherwise noted. The initial aldehyde concentration was 6.50 × 10^{-5} M in the first five runs and 6.48 × 10^{-5} M in the other runs. [b] Initial concentrations. The concentration of sodium hydroxide originally present has been corrected by about 0.2% for formation of POH^-, 0.3% for expansion on warming from room temperature to 35.2°, and 0.6% for dilution by added aldehyde solution. [c] Using a 1.00-cm cell; A_0 was 0.643 except where otherwise noted. [d] A_0 was taken to be 0.643(6.5/6.48). [e] Ionic strength 0.20 M. [f] Ionic strength 0.05 M.

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(13) The hydroxide ion concentrations used in this plot were the initial values after establishment of equilibrium between P, POH_2, and POH^- but before any cleavage occurred.

(14) The hydroxide ion concentrations plotted were the average values present during the kinetic run which were smaller than the "initial" value by as much as 0.7%.

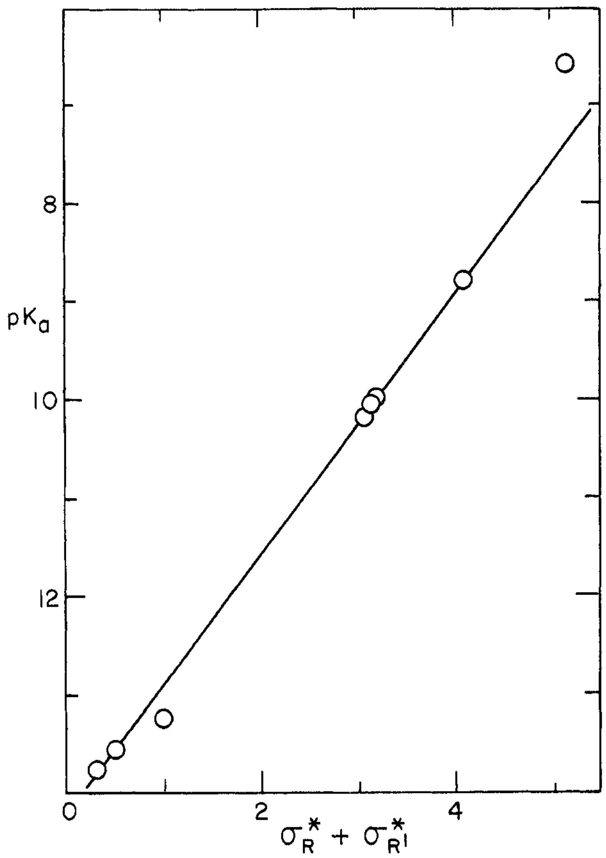

Figure 3.—Taft equation correlation of acidities of compounds of the type $RR'C(OH)_2$.

22% of the reaction took place by cleavage of the monoanion **2**.

The three runs in Table I made at hydroxide ion concentrations of 0.0116 *M* show that the reaction has a positive ionic strength effect. This is, of course, the result expected since the major part of the reaction involves the formation of a doubly charged transition state from singly charged and uncharged reactants.

Since K_E is a function of K_H and K_1, if either of the latter two constants is known the other may be calculated. We have made a rough measurement of K_H by determining the relative extinction coefficients of phenylpropargylaldehyde in 0.10 *M* aqueous sodium chloride and in 1,4-dioxane. Attributing the 29.6% decrease in the extinction coefficient at about 280 nm on going from dioxane to water entirely to hydration of the aldehyde and assuming that the aldehyde hydrate does not absorb at this wavelength gives a K_H value of 0.296. The assumption that the aldehyde hydrate does not absorb is plausible since it has the same chromophoric groups as phenylacetylene, which does not absorb near 280 nm. The K_H value obtained is uncertain, however, because of the implicit assumption that the true extinction coefficient of the aldehyde is the same in water and dioxane, in which the absorption maxima have somewhat different shapes and are at 285.5 and 278 nm, respectively. From this K_H value and the K_E obtained from spectral measurements, a value of 158 M^{-1} may be calculated for K_1 at 35° and ionic strength 0.10 *M*. According to the Debye–Hückel limiting law, this value should be the same at infinite dilution. Combination with pK_w gives a pK_a value of 11.48 for the first ionization constant of phenylpropargylaldehyde hydrate at 35° and, if K_1 is assumed to have the same value at 25°, a pK_a value of 11.80 at 25°. Available data on the ionization constants of the hydrates of aldehydes and ketones in water at 25° are listed in Table II. A plot of the pK_a values for $RR'C(OH)_2$ compounds *vs.* $\sigma^*_R + \sigma^*_{R'}$ (Figure 3) gives a straight line from which the points for the hydrates of formaldehyde and hexafluoroacetone deviate somewhat.[15] A least-squares treatment, neglecting the two deviant points, gave eq 5 for this line. Bell obtained a similar value

$$pK_a = 14.19 - 1.315\ (\sigma^*_R + \sigma^*_{R'}) \qquad (5)$$

(1.4) for ρ^* using data on four compounds.[17] From eq 5 a p*K* value of 11.77 may be calculated for phenylpropargylaldehyde hydrate, in reasonable agreement with the value we have obtained.

TABLE II

IONIZATION CONSTANTS OF HYDRATES OF ALDEHYDES AND KETONES IN WATER AT 25°

Compd	p*K*
$(CH_3)_2CHCH(OH)_2$	13.77[a]
$CH_3CH(OH)_2$	13.57[a]
$CH_2(OH)_2$	13.27[a]
$CF_3CH(OH)_2$	10.20[b]
$CCl_3CH(OH)_2$	10.04[a]
$C_6H_5C(OH)_2CF_3$	10.00[a]
$(F_2CH)_2C(OH)_2$	8.79[c]
$(CF_3)_2C(OH)_2$	6.58[c]

[a] Reference 17. [b] R. Stewart and M. M. Mocek, *Can. J. Chem.*, **41**, 1160 (1963). [c] W. J. Middleton and R. V. Lindsey, Jr., *J. Amer. Chem. Soc.*, **86**, 4948 (1964).

An approximate value for k_2 may be obtained by estimating K_2. From the value determined for pK_1, an acid-weakening factor of $10^{4.4}$ for a negative charge two atoms away calculated by the method of Branch and Calvin,[18] and a statistical factor of four, a pK_2 value of 2.8 may be calculated. This gives values of $1.58 \times 10^{-3}\ M^{-1}$ and 190 sec^{-1} for K_2 and k_2, respectively.

As an alternative to part of mechanism 1, it might be suggested that the part of the reaction involving two hydroxide ions has as its rate-controlling step attack of hydroxide ions on POH^-, either to give PO^{2-}, which is cleaved faster than it is reprotonated, or to give formate ions and phenylacetylide anions in a concerted β elimination. Such mechanisms give a rate equation identical with eq 2, except the term k_2K_2 is replaced by k_2', the rate constant for attack of hydroxide ions on POH^-. Hence, if either of these alternative mechanisms is correct, k_2' has the value 0.299 $M^{-1}\ sec^{-1}$. Since the equilibrium governed by K_2 is a proton transfer between oxygen atoms in which the equilibrium lies to the left, the work of Eigen and others[19] shows that the reverse rate constant must

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(15) The value of $\sigma^*_{CF_3}$ used in this plot was $2.8\sigma^*_{CF_3CH_2}$, or 2.58. The other values were from Taft's collection.[16]

(16) R. W. Taft, Jr., "Steric Effects in Organic Chemistry," M. S. Newman, Ed., Wiley, New York, N. Y., 1956, Chapter 13.

(17) R. P. Bell, *Advan. Phys. Org. Chem.*, **4**, 1 (1966).

(18) G. E. K. Branch and M. Calvin, "The Theory of Organic Chemistry," Prentice-Hall, Englewood Cliffs, N. J., 1941, p 204.

(19) M. Eigen, *Angew. Chem., Int. Ed. Engl.*, **3**, 1 (1964).

be about 10^{10} M^{-1} sec^{-1}. Therefore, the rate constant for the simple deprotonation of POH$^-$ by hydroxide ions must be about $10^{8.9}$ M^{-1} sec^{-1}. This exceeds the value 0.299 M^{-1} sec^{-1} by an amount that is much larger than the uncertainty of the estimate. Hence the simple deprotonation of POH$^-$ by hydroxide ions cannot be the rate-controlling step of that part of the reaction involving two hydroxide ions.

Experimental Section

Phenylpropargylaldehyde Solutions.—In order to prevent the formation of the yellow-to-red color that occurs with neat aldehyde at room temperature, the phenylpropargylaldehyde was distilled *in vacuo* into a receiver at −78° just before preparing stock solutions. About 0.01 *M* standard solutions of aldehyde in freshly purified 1,4-dioxane were prepared gravimetrically shortly after the aldehyde warmed to room temperature. All the operations were carried out under nitrogen. No decomposition of the aldehyde in the dioxane solutions could be detected over a period of several days. In 99% water–1% dioxane the aldehyde was found to have absorption maxima at 285.5, 247.5, and 236.5 nm with extinction coefficients of 10,040, 9250, and 9530 M^{-1} cm^{-1}, respectively.

Kinetic Runs.—In a typical run, 3 ml of a standard aqueous solution of sodium hydroxide containing enough sodium chloride to give an ionic strength of 0.10 *M* was pipetted into both the sample and reference cells of a Cary spectrophotometer, Model 14. After the solutions had reached thermal equilibrium at 35.2 ± 0.3°, 14 μl (20 μl in some runs) of a standard solution of phenylpropargylaldehyde in dioxane was added at a recorded time. The cell was shaken and absorbance measurements at 288.5 nm were begun within ~30 sec. In a reaction that had proceeded to about 11 half-lives (about 99.95% reaction), the absorbance was found to be 0.023. This value was taken as A_∞ for all the runs, and rate constants were calculated from the least-squares lines through plots of log $(A_t - A_\infty)$ *vs.* time. The slowest run was followed only to 24% reaction, but most of the other runs were followed to about 70% reaction. A typical kinetic plot is shown in Figure 4. The full uv spectrum of the solution that was run to 99.95% reaction showed maxima at 2343 and 2447 Å, with absorbances of 0.975 and 0.862, respectively, and shapes identical with those of authentic phenylacetylene, whose extinction coefficients were found to be 15,890 and 13,690 M^{-1} cm^{-1} at the respective wavelengths. These results correspond to a 95.9 ± 1.3% yield of phenylacetylene (assuming that sodium formate and any by-products absorb negligibly).

The kinetic runs were carried out under nitrogen. Concentrations used in the calculations were corrected for expansion from room temperature to 35.2°. In calculating A_B by extrapolation to zero time, only the first few kinetic points were used.

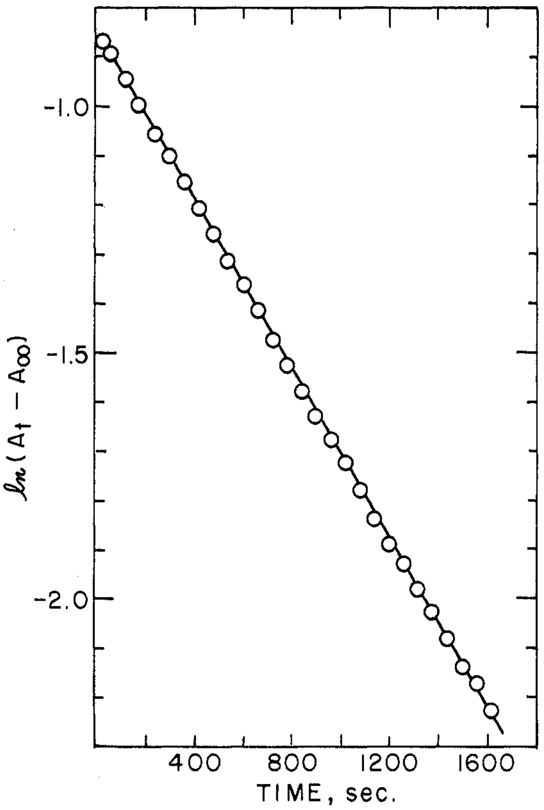

Figure 4.—Kinetic plot for cleavage of 6.48×10^{-5} *M* phenylpropargylaldehyde by 0.00932 *M* sodium hydroxide in water at 35.2°.

Rate of Hydration of Phenylpropargylaldehyde.—When 15-μl samples of 6.72×10^{-3} *M* solutions of phenylpropargylaldehyde in dioxane were injected into 3.00 ml of 0.100 *M* aqueous sodium chloride at 35° and ultraviolet measurements begun within 10 sec, the hydration reaction appeared to be essentially at equilibrium. The same technique at about 5° gave rate constants of 0.03 ± 0.01 sec^{-1} that are unreliable because of condensation on the cells and other complications that also prevented a reliable determination of the extinction coefficient of the free aldehyde by extrapolation to zero time.

Registry No.—Phenylpropargylaldehyde, 2579-22-8; sodium hydroxide, 1310-73-2.